

# CHM 129

## Problem Set #12

①

① (a)  $q = +850 \text{ J}$        $w = -382 \text{ J}$

$$\Delta E = q + w = 850 \text{ J} + (-382 \text{ J}) = 468 \text{ J}$$

Endothermic process because heat ( $q$ ) is absorbed (+).

(b)  $q = -255 \text{ J}$        $w = -P\Delta V = -(1.1 \text{ atm})(3.5 \text{ L} - 0.5 \text{ L})$

$$= -3.3 \text{ L}\cdot\text{atm} \left( \frac{101.3 \text{ J}}{1 \text{ L}\cdot\text{atm}} \right) = -334 \text{ J}$$

$$\Delta E = q + w = -255 \text{ J} + (-334 \text{ J}) = -589 \text{ J}$$

Exothermic process because heat ( $q$ ) is released (-).

(c)  $q = -6.47 \text{ kJ}$        $w = 0$

$$\Delta E = q + w = -6.47 \text{ kJ} + 0 = -6.47 \text{ kJ}$$

Exothermic process because heat ( $q$ ) is released (-).

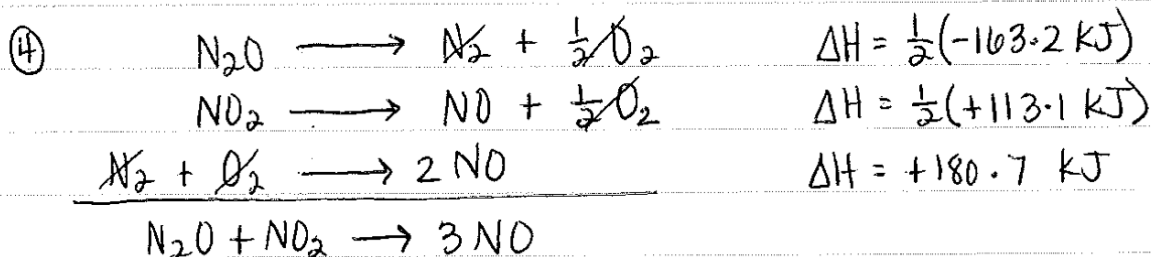
② (a) When a process occurs under constant external pressure, the enthalpy change ( $\Delta H$ ) equals the amount of heat transferred.  $\Delta H = q_p$

(b)  $\Delta H = q_p$ . If the system absorbs heat,  $q$  and  $\Delta H$  are positive and the enthalpy of the system increases.

③ (a)  $0.200 \text{ mol AgCl} \left( \frac{-65.5 \text{ kJ}}{1 \text{ mol AgCl}} \right) = -13.1 \text{ kJ}$



$$0.150 \text{ mmol} \left( \frac{10^{-3} \text{ mol}}{1 \text{ mmol}} \right) \left( \frac{+65.5 \text{ kJ}}{1 \text{ mol AgCl}} \right) = \underline{\underline{9.83 \times 10^{-3} \text{ kJ}}} = \underline{\underline{9.83 \text{ J}}}$$



$$\Delta H = \left(\frac{1}{2} \times -163.2 \text{ kJ}\right) + \left(\frac{1}{2} \times +113.1 \text{ kJ}\right) + 180.7 \text{ kJ} = +155.7 \text{ kJ}$$

$$\textcircled{5} \text{ (a) } \Delta H_R^\circ = \left[(2 \text{ mol} \times 0 \frac{\text{kJ}}{\text{mol}}) + (2 \text{ mol} \times -285.83 \frac{\text{kJ}}{\text{mol}})\right] - \left[(4 \text{ mol} \times -36.23 \frac{\text{kJ}}{\text{mol}} + (1 \text{ mol} \times 0 \frac{\text{kJ}}{\text{mol}})] = -426.74 \text{ kJ}$$

$$\text{(b) } \Delta H_R^\circ = \left[(1 \text{ mol} \times -1387.1 \frac{\text{kJ}}{\text{mol}}) + (1 \text{ mol} \times -241.82 \frac{\text{kJ}}{\text{mol}})\right] - \left[(2 \text{ mol} \times -425.6 \frac{\text{kJ}}{\text{mol}} + (1 \text{ mol} \times -395.2 \frac{\text{kJ}}{\text{mol}})] = -382.5 \text{ kJ}$$

$$\text{(c) } \Delta H_R^\circ = \left[(1 \text{ mol} \times -139.3 \frac{\text{kJ}}{\text{mol}}) + (4 \text{ mol} \times -92.30 \frac{\text{kJ}}{\text{mol}})\right] - \left[(1 \text{ mol} \times -74.8 \frac{\text{kJ}}{\text{mol}} + (4 \text{ mol} \times 0 \frac{\text{kJ}}{\text{mol}})] = -433.7 \text{ kJ}$$

$$\text{(d) } \Delta H_R^\circ = \left[(3 \text{ mol} \times -241.82 \frac{\text{kJ}}{\text{mol}}) + (2 \text{ mol} \times -400 \frac{\text{kJ}}{\text{mol}})\right] - \left[(1 \text{ mol} \times -822.16 \frac{\text{kJ}}{\text{mol}} + (6 \text{ mol} \times -92.30 \frac{\text{kJ}}{\text{mol}})] = -150 \text{ kJ}$$

⑥ (a) Both beakers of water contain the same mass of water, so they both have the same heat capacity. Object A raises the temperature of its water more than object B, so more heat was transferred from object A than B. Since both objects were heated to the same temperature initially, object A must have absorbed more heat to reach  $100^\circ$ . The greater the heat capacity of an object, the greater the heat required to produce a given rise in temperature. Thus, object A has the greater heat capacity.

(b) Since no information about the masses of the objects is given, we cannot compare the specific heats of the objects.

$$7 \quad q = m C_s \Delta T = (62.0 \text{ g})(2.42 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}})(40.5^\circ\text{C} - 13.1^\circ\text{C})$$

$$q = 4,110 \text{ J}$$

$$8 \quad 2) \quad (a) \quad q = -m_{\text{soln}} C_s \Delta T = \Delta H \quad (\text{at constant pressure})$$

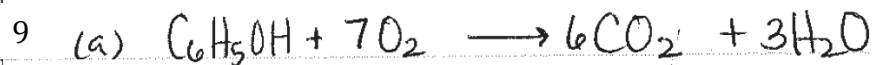
$$q = -(60.0 \text{ g} + 3.88 \text{ g})(4.184 \frac{\text{J}}{^\circ\text{C} \cdot \text{g}})(18.4^\circ\text{C} - 23.0^\circ\text{C})$$

$$q = \underline{+1,229 \text{ J}} = \underline{1,200 \text{ J}}$$

$$3.88 \text{ g} \left( \frac{1 \text{ mol}}{180.06 \text{ g}} \right) =$$

$$\Delta H = \frac{1,229 \text{ J}}{0.0484 \text{ mol}} \left( \frac{1 \text{ kJ}}{1000 \text{ J}} \right) = \underline{\underline{25 \text{ kJ/mol}}}$$

(b)  $\Delta H$  is positive  $\Rightarrow$  heat is absorbed Endothermic Process



$$(b) \quad q_{\text{R}} = -C_{\text{cal}} \times \Delta T$$

$$= -(11.66 \text{ kJ}/^\circ\text{C})(26.37^\circ\text{C} - 21.36^\circ\text{C})$$

$$q_{\text{R}} = -58.4 \text{ kJ}$$

$$\text{Per gram} = \frac{-58.4 \text{ kJ}}{1.800 \text{ g}} = \underline{\underline{-32.5 \text{ kJ/g}}}$$

$$\text{Per mole} : \frac{-58.4 \text{ kJ}}{0.01912 \text{ mol}} = \underline{\underline{-3,050 \text{ kJ/mol}}}$$

$$\text{mole} = 1.800 \text{ g} \left( \frac{1 \text{ mol}}{94.12 \text{ g}} \right) = 0.01912 \text{ mol}$$